

## Atomic Spectroscopy

### Introduction

Most substances will emit light energy if heated to a high enough temperature. For example, a fireplace poker will glow red if left in the fireplace flame for several minutes. Similarly, neon gas will glow with a bright red color when excited with a sufficiently high electrical voltage. When energy is applied to a substance, the atoms present in the substance may absorb some of that energy. Electrons within the atoms move from their normal positions to positions of higher potential energy farther away from the nucleus of the atom. Later, atoms, which have been excited by the application and absorption of energy, will "relax" and emit the excess energy they had gained. When atoms re-emit energy, often a portion of this energy is visible as light. When an electron moves from one of the fixed "orbits" to another "orbit", the attractive force of the nucleus changes by a definite amount that corresponds to a specific change in energy. The quantity of energy absorbed or emitted by an electron in moving from one allowed orbit to another is called a quantum (photon), and the energy of the particular quantum is indicated by the wavelength (or frequency) of the light emitted or absorbed by the atom. The energy of a photon is given by the Planck equation:

$$\Delta E = h\nu = hc/\lambda$$

where  $\nu$  is the frequency of light emitted or absorbed,  $\lambda$  is the wavelength corresponding to that frequency, and  $h$  is Planck's constant ( $6.626 \times 10^{-34}$  Js).

The atoms of a particular element do not emit a continuous spectrum, like the sun, but rather emit visible radiation at only certain discrete, well defined, fixed wavelengths. If the light being emitted by a particular element is passed through a prism or diffraction grating, and viewed with a spectroscope, only certain sharp bright-colored lines are seen in the resulting spectrum. The positions of these lines occur in the corresponding location (wavelength region) as in the spectrum of white light.

Line spectra for elements would later be used to refine the model of the atom. The main question to be answered was why atoms emitted light of only certain frequencies instead of a continuous spectrum like that of an incandescent light bulb. Niels Bohr answered this question and introduced the ideas of quantization and energy levels to the atomic model.

Bohr postulated that the energy of an electron when it is in a particular orbit was given by the formula:

$$E_n = -(\text{constant})/n^2$$

where  $n$  is the number of the orbit as counted out from the nucleus ( $n = 1$  means the first orbit,  $n = 2$  means the second orbit, etc.) and is called the principle quantum number. The proportionality constant in Bohr's theory is called the Rydberg constant ( $R_H$ ) and has the value  $2.18 \times 10^{-18}$  J. According Bohr's theory, if an electron were to move from an outer "orbit" ( $n_{\text{outer}}$ ) to an inner "orbit" ( $n_{\text{inner}}$ ), a photon of light should be emitted, having the energy given by

$$\Delta E = E_{\text{inner}} - E_{\text{outer}} = -R_H[(1/n_{\text{inner}}^2) - (1/n_{\text{outer}}^2)]$$

The wavelength ( $\lambda$ ) of this photon would be given by the Planck formula as

$$\lambda = hc/\Delta E$$

Bohr performed calculations of wavelengths for various values of the principal quantum number,  $n$ , and found that the predicted wavelengths from the theory agreed exactly with experimental wavelengths measured with a spectroscope. Bohr even went so far as to predict emissions by hydrogen atoms in other regions of the EM spectrum (UV, IR) that had not been observed experimentally. His predictions were confirmed almost immediately. These series of non-visible spectra are known as the Lyman and Paschen series, respectively.

Bohr's simple atomic theory of an electron moving between fixed orbits helped to greatly explain observed spectra and formed the basis for the detailed atomic theory for more complex atoms with more than one electron. The spectra of larger atoms are considerably more complicated than that of hydrogen, but generally a *characteristic* spectrum can be seen. Bohr's theory for hydrogen accounted on a microscopic basis for the macroscopic phenomena of spectral emission lines.

The light emitted from an element may be analyzed using a spectroscope. A spectroscope includes four major features: a slit for admitting a narrow, collimated beam of light; a prism or diffraction grating that spreads the incident light into its component wavelengths; a telescope for viewing the spectrum; and an illuminated reference scale to determine the wavelength of each bright line.

In this experiment, you will use the equations developed by Balmer, Rydberg, and Bohr to predict the wavelengths and energies of light emitted for transitions in the hydrogen atom. These values will be compared with your experimental values and the accepted values found in your text.

### Equipment/Materials

hydrogen discharge tube

spectroscope

### Procedure

#### *Part 1 Hydrogen Spectrum*

1. Obtain a spectroscope. Look through the eyepiece, and make sure that the hydrogen lamp can be viewed through the slit when the lamp is turned on.
2. When everyone is ready, the instructor will turn on the hydrogen lamp. Look through the eyepiece and adjust the slit opening of the spectroscope so that the spectral lines are as bright and sharp as possible in the illuminated range of the spectroscope (to the side of the slit).
3. Record the color and wavelength for each line in the visible spectrum of hydrogen (Balmer Series). Be sure to use both the handheld spectroscope and the larger spectroscope to record wavelength.

### Calculations

1. Using Bohr's equation, calculate the energy that would be predicted for each transition in the Balmer series. All transitions are from higher energy levels to the  $n = 2$  energy level. For the visible spectrum, the maximum energy level is  $n = 6$ . For these predicted energies, calculate the wavelength of light for each emitted photon.
2. Using the wavelengths of light from hydrogen spectrum you recorded with the spectroscope, calculate the energies of these bright lines.
3. Look up the actual value for the wavelengths of light in the Balmer series of the hydrogen spectrum. Calculate the energies for the actual wavelengths.
4. Determine the percent error between your predicted and measured values from the actual values for these energies.

### Questions and Conclusions

1. Were your predicted and measured values for the energy emitted close to the actual values of energy released in hydrogen's spectrum? Explain why or why not.
2. Why did you calculate the energies ranging only from  $n = 6$  to  $n = 2$ ?
3. The Paschen series lines in the atomic spectrum of hydrogen result from transitions from  $n > 3$  to  $n = 3$ . Calculate the wavelength in nm of a line in this series resulting from an  $n = 6$  to  $n = 3$  transition. Where would this line be found in the electromagnetic spectrum?
4. Calculate the wavelength of light emitted when an electron changes from  $n = 3$  to  $n = 1$  in the hydrogen atom. In what region of the spectrum is this radiation found?